ARTHAM
RESOURCE MATERIAL
NO. 1 EDUCATIONAL RESOURCES

ARTHAM
RESOURCE MATERIAL

SCHOOL OF EDUCATORS
SOE
Learn | Unlearn | Relearn

PREVIOUS YEAR QUESTION PAPERS

WITH SOLUTIONS

CLASS 12

CHEMISTRY

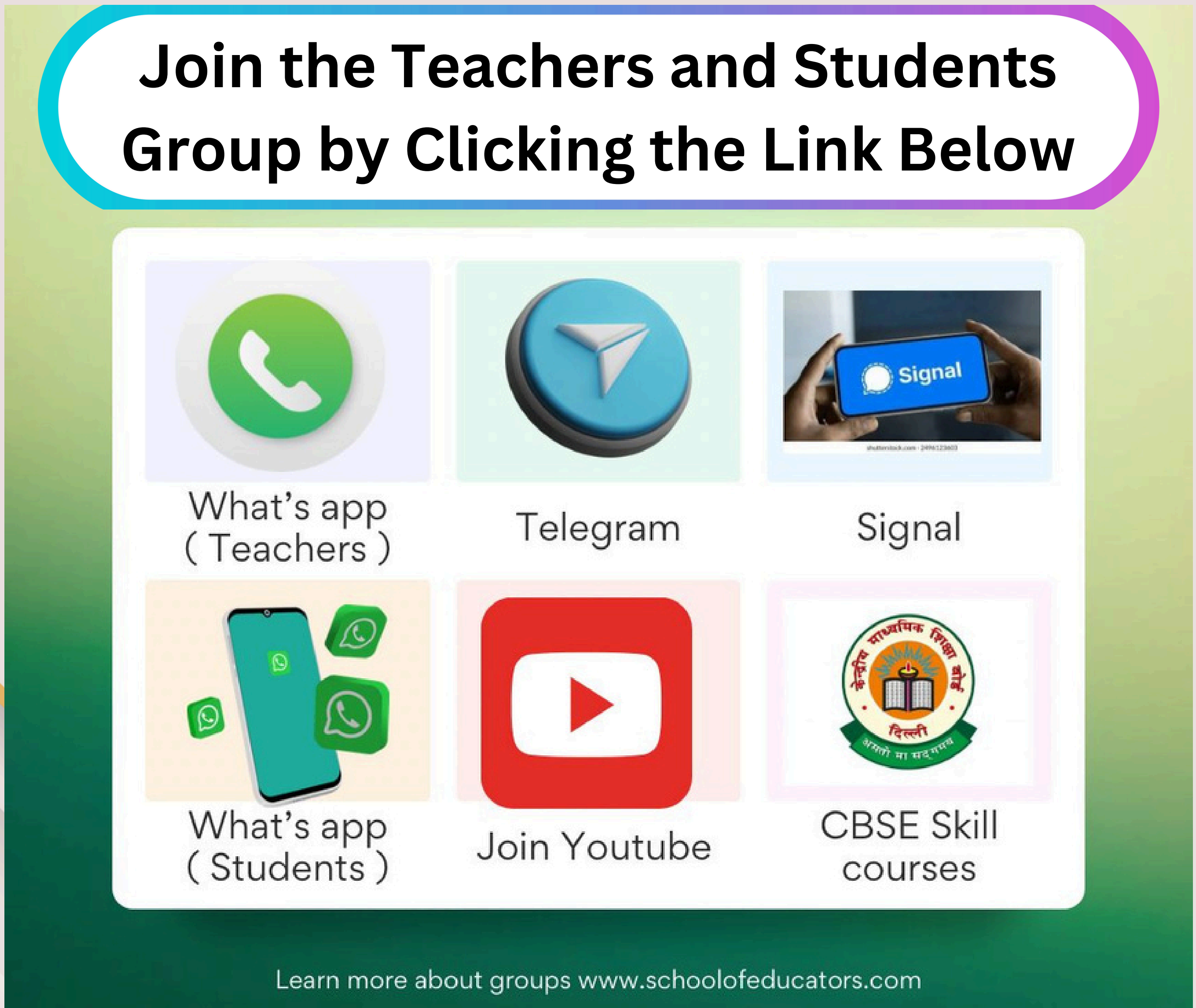
CHAPTER WISE
TOPIC WISE
SOLVED PAPERS
From 2014 to 2024

Dear Teachers and Students,

Join School of Educators' exclusive WhatsApp, Telegram, and Signal groups for FREE access to a vast range of educational resources designed to help you achieve 100/100 in exams! Separate groups for teachers and students are available, packed with valuable content to boost your performance.

Additionally, benefit from expert tips, practical advice, and study hacks designed to enhance performance in both CBSE exams and competitive entrance tests.

Don't miss out—join today and take the first step toward academic excellence!



Class 12 Chemistry

Previous Year Questions

Chapter-2 : Electrochemistry

1. Electrochemical Cells And Electrode Potential

MCQ

1. Assertion (A): For a Daniell cell, $Zn/Zn^{2+}(1M) || Cu^{2+}(1M)/Cu$ with

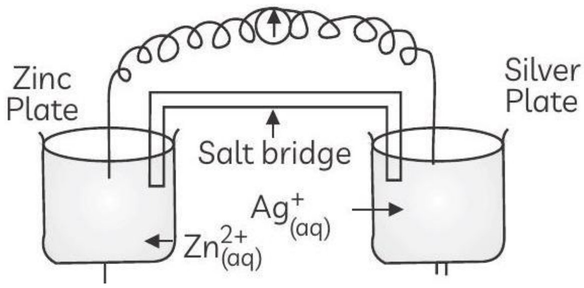
E°_{cell} =1.1 V, if the external opposing potential is more than 1.1 V, the electrons flow from Cu to Zn.
Reason (R): Cell acts like a galvanic cell.
(A) Both Assertion (A) and Reason (R) are correct, and Reason (R) is the correct explanation for Assertion (A).
(B) Both Assertion (A) and Reason (R) are correct, but Reason (R) is not the correct explanation for Assertion (A).
(C) Assertion (A) is correct, but Reason (R) is incorrect.
(D) Assertion (A) is incorrect, but Reason (R) is correct. (2024)

Case Based Qs (4-5 mark)

Read the following passages and answer the questions that follow:

2. Oxidation-reduction reactions are commonly known as redox reactions.

They involve transfer of electrons from one species to another. In a spontaneous reaction, energy is released which can be used to do useful work. The reaction is split into two half reactions. Two different containers are used and a wire is used to drive the electrons from one side to the other and a Voltaic/Galvanic cell is created. It is an electrochemical cell that uses spontaneous redox reactions to generate electricity. A salt bridge also connects to the half cells. The reading of the voltmeter gives the cell voltage or cell potential or electromotive force. If E^{o}_{cell} is positive, the reaction is spontaneous and if it is negative, the reaction is non-spontaneous and is referred to as electrolytic cell. Electrolysis refers to the decomposition of a substance by an electric current. One mole of electric charge when passed through a cell it will discharge half a mole of a divalent metal ion such as Cu^{2+}. This was first formulated by Faraday in the form of laws of electrolysis. The conductance of material is the property of materials due to which a material allows the flow of ions through itself and thus conducts electricity. Conductivity is represented by κ and it depends upon nature and concentration of electrolyte temperature etc. A more common term molar conductivity of a solution at a given concentration is conductance of the volume of solution containing one mole of electrolyte kept between two electrodes, with the unit area of cross-section and distance of unit length. Limiting molar conductivity of weak electrolytes cannot be obtained graphically.



(A) Is silver plate act as anode or cathode?
(B) What will happen if the salt bridge is removed?
(C) When does electrochemical cell behaves like an electrolytic cell?
(D) (i) What will happen to the concentration of Zn^{2+} and $Ag^{+}$when $E_{cell} = 0$
(ii) Why does conductivity of a solution decreases with dilution?

OR

The molar conductivity of a $1.5M$ solution of an electrolyte is found to be $138.9 scm^2 mol^{-1}$. Calculate the conductivity of this solution. (Term-2 2020)

VSA (1-3 mark)

3. Write electrode reactions and calculate the E_{cell} of the following cell at $298K$:
$Ni_{(s)} | Ni^{2+}(0.001M) \| Ag^{+}(0.1M) | Ag_{(s)}$
[Given: E° for $Ni^{2+} | Ni = 0.25V, Ag^{+} | Ag = 0.80V$] (Term-2 2020)

4. Represent the cell in which the following reaction takes place. The value of $E°$ for the cell is $1.260V$. What is the value of E_{cell} ? (Term-2 2020)
$2Al_{(s)} + 3Cd^{2+}{}_{(0.1M)} \rightarrow 3Cd_{(s)} + 2Al^{3+}(0.01M)$

5. Give two points of difference between the electrolytic cell and electrochemical cell. (2020)

6. Calculate $\Delta G°$ for the reaction.

$$Zn_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Zn^{2+}{}_{(aq)} + Cu_{(s)}$$

Given: E° for $Zn^{2+}/Zn = -0.76V$ and
E° for $Cu^{2+}/Cu = +0.34V$
$R = 8.314 JK^{-1}mol^{-1}$,
$F = 96500 Cmol^{-1}$,

(2020)

7. Calculate the maximum work and $\log K_c$ for the given reaction at $298K$:

$$Ni_{(s)} + 2Ag^{+}_{(aq)} \rightarrow Ni^{2+}{}_{(aq)} + 2Ag_{(s)}$$

Given:

$$E^{\circ}_{Ni^{2}+}/Ni = -0.25V E^{\circ}_{Ag^{+}/Ag} = +0.80V$$

$$1F = 96500 Cmol^{-1}$$

(2020)

8. Calculate e.m.f. of the following cell:

$Zn_{(s)}|Zn^{2+}(0.1M)||Ag^{+}(0.01M)|Ag_{(s)}$

Given: E° for $Zn^{2+}/Zn = -0.76V, E^{\circ}$ for $Ag^{+}/Ag = +0.80V$ [Given: $\log 10 = 1$] (2020)

9. Define electrochemical cell. What happens if external potential applied becomes greater than E°_{Cell} of electrochemical cell? (2019,2016)

10. Calculate the equilibrium constant for the reaction:

$$Cd^{2+}{}_{(aq)} + Zn_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + Cd_{(s)}$$

If $E^{\circ}_{Cd^{2+}/Cd} = -0.403V; E^{\circ}_{Zn^{2+}/Zn} = -0.763V$ (2019)

11. E°_{cell} for the given redox reaction is $2.71V$

$Mg_{(s)} + Cu^{2+}(0.01M) \rightarrow Mg^{2+}(0.001M) + Cu_{(s)}$

Calculate E_{cell} for the reaction. Write the direction of flow of current when an external opposite potential applied is:

(i) less than $2.71V$ and

(ii) greater than $2.71V$ (2019)

12. Write the cell reaction and calculate the e.m.f. of the following cell at $298K$:

$Sn_{(s)}|Sn^{2+}(0.004M) \parallel H^{+}(0.020M)|H_{2(g)}$ (1 bar) $|Pt_{(s)}$

(Given, $E^{\circ}_{Sn^{+2}/Sn} = -0.14V$) (2018)

13. For the reaction.

$2AgCl_{(s)} + H_{2(g)}(1atm) \rightarrow 2Ag_{(s)} + 2H^{+}(0.1M) + 2Cl^{-}(0.1M), \Delta G^{\circ} = -43600J$ at $25^{\circ}C$. Calculate the e.m.f of the cell. [$[\log 10^{-n} = -n]$ (2018)

14. Calculate e.m.f of the following cell at 298K:

$2Cr_{(s)} + 3Fe^{2+}(0.1M) \rightarrow 2Cr^{3+}(0.01M) + 3Fe_{(s)}$

Given: $E_{Cr^{3+}/Cr} = -0.74V E_{Fe^{2+}/Fe} = -0.44V$ (2016)

15. Calculate E_{cell} for the following reaction at 298K.

$2Al_{(s)} + 3Cu^{+2}(0.01M) \rightarrow 2Al^{+3}(0.01M) + 3Cu_{(s)}$; Given: $E^\circ_{\text{cell}} = 1.98V$

(2016)

16. Calculate $\Delta_r G^\circ$ for the reaction: $Mg_{(s)} + Cu^{2+}_{(aq)} \rightarrow Mg^{2+}_{(aq)} + Cu_{(s)}$

Given: $E^\circ_{\text{cell}} = +2.71V, 1F = 96500 C mol^{-1}$ (2014)

17. Calculate emf of the following cell at 298K: $Mg_{(s)} | Mg^{2+}(0.1M) \parallel$

$Cu^{2+}(0.01M) | Cu_{(s)}$ [Given $E^\circ_{\text{cell}} = +2.71V$, 1 Faraday $= 96500 C mol^{-1}$].

(2014)

18. Estimate the minimum potential difference needed to reduce Al_2O_3 at $500^\circ C$. The free energy change for the decomposition reaction:

$\frac{2}{3}Al_2O_3 \rightarrow \frac{4}{3}Al + O_2$ is $\Delta G = +960 kJ$, $F = 96500 C mol^{-1}$ (2014)

19. Equilibrium constant (K_c) for the given reaction is 10 . Calculate E°_{cell}.

$A_{(s)} + B^{2+}_{(aq)} \rightleftharpoons A^{2+}_{(aq)} + B_{(s)}$ (2014)

20. Calculate the emf of the following cell:

$Ni(s) + 2Ag^+(0.01\ M) \longrightarrow Ni^{2+}(0.1\ M) + 2Ag(s)$

Given that $E^\circ_{\text{cell}} = 1.05\ V$, $\log 10 = 1$

(2024)

2. Conductivity Of Electrolytic Solutions

MCQ

21. Dilution affects both conductivity as well as molar conductivity. Effect of dilution on both is as follows:

(A) both increase with dilution.

(B) both decrease with dilution.

(C) conductivity increases whereas molar conductivity decreases on dilution.

(D) conductivity decreases whereas molar conductivity increases on dilution.

(2024)

22. Which of the following solutions will have the highest conductivity 298K ?

(a) $0.01 M HCl$ solution

(b) $0.1 M HCl$ solution

(c) $0.01 M CH_3COOH$ solution

(d) $0.1 M CH_3COOH$ solution (2023)

23. Which of the following solutions of KCl will have the highest value of molar conductivity?

(a) 0.01M

(b) 1M

(c) 0.5 M

(d) 0.1 M (2023)

In the following question (Q. No. 24-26) a statement of assertion (A) followed by a statement of reason (R) is given. Choose the correct answer out of the following choices.

(a) Both (A) and (R) are true and (R) is the correct explanation of (A).

(b) Both (A) and (R) are true but (R) is not the correct explanation of (A).

(c) (A) is true, but (R) is false.

(d) (A) is false, but (R) is true.

24. Assertion (A): Conductivity decreases with decrease in concentration of electrolyte.

Reason (R): Number of ions per unit volume that carry the current in a solution decreases on dilution. (2023)

25. Which of the following option will be the limiting molar conductivity of CH_3COOH if the limiting molar conductivity of CH_3COONa is $91 Scm^2 mol^{-1}$? Limiting molar conductivity for individual ions are given in the following table.

S. No.	Ions	Limiting molar conductivity/Scm^2 mol^{-1}
(1)	H^+	349.6
(2)	Na^+	50.1
(3)	K^+	73.5
(4)	OH^-	199.1

(a) $350 scm^2 mol^{-1}$

(b) $375.3 scm^2 mol^{-1}$

(c) $390.5 scm^2 mol^{-1}$

(d) $340.4 scm^2 mol^{-1}$ (2020)

26. The molar ionic conductivities of Ca^{2+} and Cl^- are 119.0 and 76.3 S cm^2 mol^{-1} respectively. The value of limiting molar conductivity of $CaCl_2$ will be:

A) 195.3 S cm^2 mol^{-1}

B) 43.3 S cm^2 mol^{-1}

C) 314.3 S cm^2 mol^{-1}

D) 271.6 S cm^2 mol^{-1}

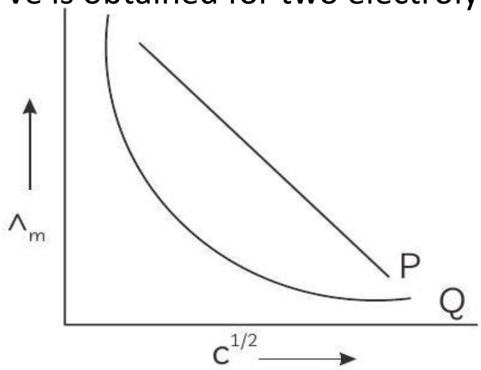
(2024)

VSA (1-3 mark)

27. Solutions of two electrolytes 'A' and 'B' are diluted. The Λ_m of 'B' increases 1.5 times while that of 'A' increases 25 times. Which of the two is a strong electrolyte? Justify your answer. Graphically show the behavior of ' A ' and 'B'.

(Term-2 2022)

28. In the plot of molar conductivity (Λ_m) vs square root of concentration $(c^{1/2})$ following curve is obtained for two electrolytes A and B :

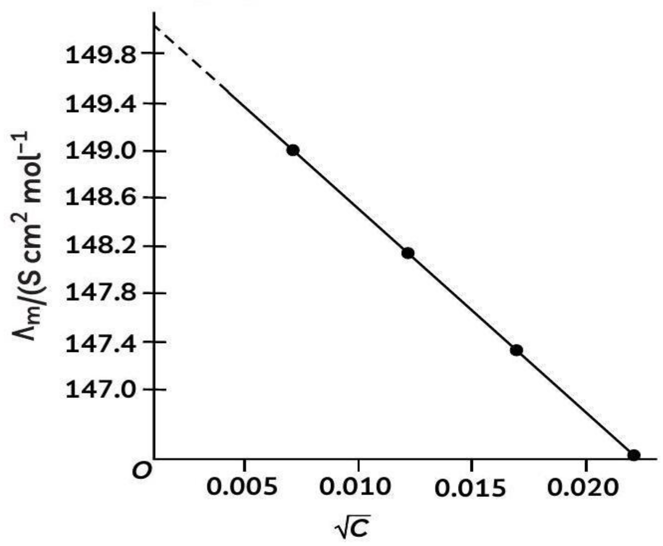


Answer the following:
(i) Predict the nature of electrolytes A and B.
(ii) What happens on extrapolation of Λ_m to concentration approaching zero for electrolytes A and B? (2019)

29. "Conductivity is a measurement of the ability of an aqueous solution to transfer an electrical current. The current is carried by ions, and therefore, the conductivity increases with the concentration of ions present in solution, their mobility, and temperature of the water. Conductivity measurements are related to ionic strength." It is observed that on dilution, conductivity of CH_3COOH decreases. Support this statement by giving proper explanation.
(2018)

30. Give reason:
Conductivity of CH_3COOH decreases on dilution. (2018)

31. Consider the graph:



Observe the graph and identify and define the chemical term when concentration approaches zero. Why conductivity of an electrolyte solution decreases with the decrease in concentration? (2015)

32. Define the term degree of dissociation. Write an expression that relates the molar conductivity of a weak electrolyte to its degree of dissociation. (2015)

33. Define conductivity and molar conductivity for the solution of an electrolyte. Discuss their variation with concentration. (2015)

34. Define the term degree of dissociation. Write an expression that relates the molar conductivity of a weak electrolyte to its degree of dissociation.

35. State Kohlrausch law of independent migration of ions. Why does the conductivity of a solution decrease with dilution? (2014)

36. Define the terms conductivity and molar conductivity for the solution of an electrolyte. Comment on their variation with temperature. (2014)

37. **Question:** The conductivity of 0.2 M solution of KCl is $2.48 \times 10^{-2}\ \mathrm{S\ cm^{-1}}$. Calculate conductivity and degree of dissociation (α). Given:

$\lambda_0^{K^+} = 73.5\ \mathrm{S\ cm^2 mol^{-1}}$

$\lambda_0^{Cl^-} = 76.5\ \mathrm{S\ cm^2 mol^{-1}}$

(2024)

38. Resistance of a conductivitycell filled with0.2 molL 1KCl solution is 200n. If the resistance of the same cell when filled with 0.05 molL-1KCl solution is 620n, calculate the conductivity and molar conductivity of 0.05molL-1KCl solution.Theconductivityof0.2molL-1KCl solution is 0.0248Scm-1. (2024)

Long (4-5 mark)

39. (A) Can we construct an electrochemical cell with two half-cells composed of $ZnSO_4$ solution and zinc electrodes? Explain your answer.
(B) Calculate the λ_m° for Cl^-ion from the data given below: $\lambda_{\mathrm{m}}^{\circ}\mathrm{MgCl}_2 = 258.6\mathrm{scm}^2\mathrm{mol}^{-1}$ and $\lambda_m^0\mathrm{Mg}^{2+} = 106\mathrm{Scm}^2\mathrm{mol}^{-1}$
(C) The cell constant of a conductivity cell is $0.146\mathrm{cm}^{-1}$. What is the conductivity of $0.01\mathrm{M}$ solution of an electrolyte at $298\mathrm{K}$, if the resistance of the cell is 1000 Ohm? (2023)

40. (A) Conductivity of $2 \times 10^{-3}\mathrm{M}$ methanoic acid is $8 \times 10^{-5}\mathrm{Scm}^{-1}$. Calculate its molar conductivity and degree of dissociation if $\Lambda^{\circ}{}_{\mathrm{m}}$ for methanoic acid is $404\mathrm{Scm}^2\mathrm{mol}^{-1}$.
(B) Calculate the $\Delta_r G$ and $\log K_c$, for the given reaction at $298\mathrm{K}$:
$\mathrm{Ni}_{(s)} + 2\mathrm{Ag}^+_{(aq)} \rightleftharpoons \mathrm{Ni}^{2+}{}_{(aq)} + 2\mathrm{Ag}_{(s)}$
Given: $\mathrm{E}^{\circ}_{\mathrm{Ni}^{+2}/\mathrm{Ni}} =- 0.25\mathrm{V}, \mathrm{E}^{\circ}_{\mathrm{Ag}^+/\mathrm{Ag}} =+ 0.80\mathrm{V}$
$1\mathrm{F} = 96500\mathrm{Cmol}^{-1}$ (2023)

41. (A) Why does the cell voltage of a mercury cell remain constant during its lifetime?
(B) Write the reaction occurring at anode and cathode and the products of electrolysis of aq. KCl.
(C) What is the pH of HCl solution when the hydrogen gas electrode shows a potential of $-0.59\mathrm{V}$ at standard temperature and pressure? (2022)

42. (A) Molar conductivity of substance ' A ' is $5.9 \times 10^3\mathrm{S/m}$ and ' B ' is $1 \times 10^{-16}\mathrm{s/m}$. Which of the two is most likely to be copper metal and why?
(B) What is the quantity of electricity in Coulombs required to produce 4.8g Mg from molten MgCl_2 ? How much Ca will be produced if the same amou of electricity was passed through molten CaCl_2 ? (Atomic mass of $\mathrm{Mg} = 24$ atomic mass of $\mathrm{Ca} = 40\mathrm{u}$).
(C) What is the standard free energy change for the following reaction at room temperature? Is the reaction spontaneous?
$\mathrm{Sn}_{(s)} + 2\mathrm{Cu}^{2+}{}_{(aq)} \rightarrow \mathrm{Sn}^{2+}{}_{(aq)} 2\mathrm{Cu}_{(s)}$ (202

43. Molar conductivity of acetic acid solution is $39.0 \mathrm{Scm}^2\mathrm{mol}^{-1}$. If limiting molar conductivities of NaCl, HCl and CH_3COONa are 126.4, 425.9 and $91.0 \mathrm{scm}^2\mathrm{mol}^{-1}$ respectively.

(A) Calculate limiting molar conductivity of acetic acid.

(B) How much acetic acid is present in unionized form for given solution?

(Term-2 2022)

44. (A) Calculate the degree of dissociation of $0.0024\mathrm{M}$ acetic acid if conductivity of this solution is $8.0 \times 10^{-5} \mathrm{Scm}^{-1}$.

$$\lambda^\circ_{H^+} = 349.6 \mathrm{Scm}^2\mathrm{mol}^{-1};$$

$$\lambda^\circ_{CH_3CoO^-} = 40.9 \mathrm{scm}^{-1}$$

(B) Solutions of two electrolytes 'A' and ' B ' are diluted. The limiting molar conductivity of 'B' increases to a smaller extent while that of ' A ' increases to a much larger extent comparatively. Which of the two is a strong electrolyte?

(2019)

45. Calculate the degree of dissociation (α) of acetic acid if its molar conductivity is 39.05 $\mathrm{Scm}^2\mathrm{mol}^{-1}$.

Given: $\lambda^0_{H^+} = 349.6 \mathrm{Scm}^2\mathrm{mol}^{-1}$ and $\lambda^\circ_{CH_3COO^-} = 40.9 \mathrm{Scm}^2\mathrm{mol}^{-1}$.

(2017)

Numerical Type Qs (1 - 3 marks)

46. The conductivity of $0.20 \mathrm{molL}^{-1}$ solution of KCl is $2.48 \times 10^{-2} \mathrm{Scm}^{-1}$.

Calculate its molar conductivity and degree of dissociation (α).

Given: $\lambda^0_{(K^+)} = 73.5 \mathrm{Scm}^2\mathrm{mol}^{-1}$ and $\lambda^\circ_{(Cl^-)} = 76.5 \mathrm{Scm}^2\mathrm{mol}^{-1}$ · (2015)

47. Resistance of a conductivity cell filled with $0.1 \mathrm{molL}^{-1}$ KCl solution is 100Ω. If the resistance of the same cell when filled with $0.02 \mathrm{molL}^{-1}$ KCl solution is 520Ω, calculate the conductivity and molar conductivity of $0.02 \mathrm{molL}^{-1}$ KCl solution is $1.29 \times 10^{-2} \Omega^{-1}\mathrm{cm}^{-1}$. (2014)

4. Electrolytic Cell And Electrolysis

MCQ

48. Which of the following cells was used in the Apollo space program?

(A) Mercury cell

(B) H_2–O_2 fuel cell

(C) Dry cell

(D) Ni-Cd cell

(2024)

VSA (1-3 mark)

49. How many coulombs are required for the oxidation of 1mol of H_2O_2 to O_2 ? (2020)

50. Give reason:

On the basis of E° values, O_2 gas should be liberated at anode but Cl_2 gas is liberated at anode in the electrolysis of aqueous $NaCl$. (2020)

51. How much charge in terms of Faraday is required to reduce one mole of MnO_4^- to Mn^{2+} ? (2020)

52. How many coulombs of electricity are required for the oxidation of 1 mole of H_2O to O_2 ? (2020)

53. Following reactions occur at cathode during electrolysis of aqueous silver chloride solution:

$$Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}, E^\circ = +0.80V$$

$$H^+_{(aq)} + e^- \rightarrow \frac{1}{2}H_{2(g)}, E^\circ = 0.00V$$

On the basis of standard reduction potential (E° value), which reaction is feasible at cathode and why? (2017,2015)

54. How much charge is required for the reduction of 1 mol of Zn^{2+} to Zn ? (2015)

55. Faraday's laws of electrolysis have quantitative relationships based on the electrochemical research published by Michael Faraday in 1833. State the Faraday's first Law of electrolysis. (2015)

56. Following reactions occur at cathode during the electrolysis of aqueous copper (II) chloride solution:

$$Cu^{2+}{}_{(aq)} + 2e^- \rightarrow Cu_{(s)}, E^\circ = +0.34V$$

$$H^+(aq) + e^- \rightarrow \frac{1}{2}H_{2(s)}, E^\circ = 0.00V$$

On the basis of their standard reduction electrode potential (E°) values, which reaction is feasible at the cathode and why? (2015)

57. Calculate emf of the following cell at 25°C:

$Sn/Sn^{2+}(0.001\ M)\ ||\ H^{+}(0.01\ M)|H_2(g)(1\ bar)|Pt_{(s)}$

Given: $E^{\circ}(Sn^{2+}/Sn) = -0.14\ V,\ E^{\circ}(H^{+}/H_2) = 0.00\ V$ (log 10 = 1) (2024)

Numerical Type Qs (1 - 3 marks)

58. Chromium metal is electroplated using an acidic solution containing CrO_3 according to the following equations:

$$CrO_{3(aq)} + 6H^{+} + 6e^{-} \rightarrow Cr_{(s)} + 3H_2O$$

Calculate how many grams of chromium will be electroplated by 24,000 coulombs. How long will it take to electroplate 1.5g chromium using 12.5A current?
[Atomic mass of Cr $= 52 gmol^{-1}, 1F = 96500 Cmol^{-1}$] (2019)

59. A steady current of 2 amperes was passed through two electrolytic cells X and Y connected in series containing electrolytes $FeSO_4$ and $ZnSO_4$ until 2.8g of Fe deposited at the cathode of cell X. How long did the current flow? Calculate the mass of Zn deposited at the cathode of cell Y. (Molar mass : $Fe = 56 gmol^{-1}, Zn = 65.3 gmol^{-1}$, $1F = 96500 (mol^{-1})$) (2019)

60. Calculate the mass of Ag deposited at cathode when a current of 2 ampere was passed through a solution of $AgNO_3$ for 15 minutes.
[Given: Molar mass of Ag $= 108 gmol^{-1}$, $F = 96500 Cmol^{-1}$] (2017)

61. (A) The cell in which the following reaction occurs:

$$2Fe^{3+}_{(aq)} + 2I^{-}_{(aq)} \longrightarrow 2Fe^{2+}_{(aq)} + I_{2(s)}$$

has $E^{\circ}_{cell} = 0.236 V$ at 298K. Calculate the standard Gibbs energy of the cell reaction. (Given: $1F = 96500 Cmol^{-1}$)
(B) How many electrons flow through a metallic wire if a current of $0.5A$ is passed for 2 hours? (Given: 1F = $96500 Cmol^{-1}$) (2017)

62. State Faraday's first Law of electrolysis. How much charge in terms of Faraday is required for the reduction of $1 mol$ of Cu^{2+} to Cu ? (2014)

63. A solution of $Ni(NO_3)_2$ is electrolysed between platinum electrodes using a current of 5 amperes for 20 minutes. What mass of Ni is deposited at the cathode? (2014)

4. Commercial Cells/ Batteries

Case Based Qs (4-5 mark)

64. The lead-acid battery represents the oldest rechargeable battery technology. Lead acid batteries can be found in a wide variety of applications including small-scale power storage such as UPS systems, ignition power sources for automobiles, along with large, grid-scale power systems. The spongy lead act as the anode and lead dioxide as the cathode. Aqueous sulphuric acid is used as an electrolyte. The half-reactions during discharging of lead storage cells are:

Anode: $\mathrm{Pb_{(s)} + SO_4^{2-}}\ _{(aq)} \rightarrow \mathrm{PbSO_{4(s)}} + 2e^-$

Cathode: $\mathrm{PbO_{2(s)} + 4H^+}\ _{(aq)} + \mathrm{SO_4^{2-}(aq) + 2e^-}$
$\rightarrow \mathrm{PbSO_{4(s)} + 2H_2O}$

There is no safe way of disposal and these batteries end - up in landfills. Lead and sulphuric acid are extremely hazardous and pollute soil, water as well as air. Irrespective of the environmental challenges it poses, lead acid batteries have remained an important source of energy.

Designing green and sustainable battery systems as alternatives to conventional means remains relevant. Fuel cells are seen as the future source of energy. Hydrogen is considered a green fuel. Problem with fuel cells at present is the storage of hydrogen. Currently, ammonia and methanol are being used as a source of hydrogen for fuel cell. These are obtained industrially, so add to the environmental issues.

If the problem of storage of hydrogen is overcome, is it still a "green fuel?" Despite being the most abundant element in the Universe, hydrogen does not exist on its own so needs to be extracted from the water using electrolysis or separated from carbon fossil fuels. Both of these processes require a significant amount of energy which is currently more than that gained from the hydrogen itself. In addition, this extraction typically requires the use of fossil fuels. More research is being conducted in this field to solve these problems. Despite the problem of no good means to extract Hydrogen, it is a uniquely abundant and renewable source of energy, perfect for our future zero-carbon needs.

Answer the following questions:

(A) How many coulombs have been transferred from anode to cathode in order to consume one mole of sulphuric acid during the discharging of lead storage cell?

(B) How much work can be extracted by using lead storage cell if each cell delivers about of voltage?

(C) Do you agree with the statement"Hydrogen is a green fuel."? Give your comments for and against this statement and justify your views.

OR

Imagine you are a member of an agency funding scientific research. Which of the
following projects will you fund and why?
(i) Safe recycling of lead batteries
(ii) Extraction of hydrogen (2023)

VSA (1-3 mark)

65. Give reasons:

(A) Mercury cell delivers a constant potential during its life time.
(B) In the experimental determination of electrolytic conductance, Direct Current (DC) is not used.

OR

Define fuel cell with an example. What advantages do the fuel cells have over primary and secondary batteries? (2023)

66. Corrosion is an electrochemical phenomenon. The oxygen in moist air reacts as follows:

$$O_{2(g)} + 2H_2O_{(\eta)} + 4e^- \rightarrow 4OH^-_{(aq)}$$

Write down the possible reactions for corrosion of zinc occurring at anode, cathode, and overall reaction to form a white layer of zinc hydroxide. (2022)

67. Name the type of cell which was used in Apollo space programme. (2020)

68. Give two advantages of fuel cells. (2020,2018)

69. "Fuel cells work like batteries, but they do not run down or need recharging. A fuel cell consists of a negative electrode (or anode) and a positive electrode (or cathode), sandwiched around an electrolyte." Define fuel cells by writing its working in a sentence. (2018,2017,2015,2014)

70. Define fuel cell and write its two advantages. (2018)

71. Write the name of the cell which is generally used in hearing aids. Write the reactions taking place at the anode and the cathode of this cell. (2017)

72. Write the name of the cell which is generally used in inverters. Write the reactions taking place at the anode and the cathode of this cell. (2017)

73. Using the $E°$ values of A and B, predict which is better for coating the surface of iron to prevent corrosion and why?

$$\left[E^{0}_{(Fe^{2+}/Fe)} = 0.44V\right]$$

$$E°_{(A^{2+}/A)} = -2.37V; E°_{(B^{2+}/B)} = 0.14V$$

(2016)

74. "The storage battery, secondary battery, or charge accumulator is a cell or combination of cells in which the cell reactions are reversible." Give a brief insight about the secondary batteries. (2015)

75. What type of battery is mercury cell? Why it is more advantageous than dry cell? (2015)

ARTHAM

RESOURCE MATERIAL

NO. 1 EDUCATIONAL RESOURCES

Class 12 Chemistry

PYQ (Solution)

Chapter-2 : Electrochemistry

1. Electrochemical Cells And Electrode Potential

MCQ

1. (A) Both Assertion (A) and Reason (R) are correct, and Reason (R) is the correct explanation for Assertion (A).
Explanation: Both Assertion (A) and Reason (R) are correct, and Reason (R) is the correct explanation for Assertion (A).

Case Based Qs (4-5 mark)

2. (A) Silver plate acts as cathode.
Explanation: In an electrochemical cell, reduction half-cell is on right side which acts as cathode.
(B) If salt bridge is removed, no current will flow.
Explanation: Function of salt bridge is to complete the circuit by connecting the two electrolyte solutions of the cell. If it is removed, circuit will not be completed and no current will flow.
(C) An electrochemical cell behaves as an electrolytic cell when there is an application of an external and opposite potential on the galvanic cell.
Explanation: When $E_{\text{ext}} > E_{\text{cell,}}$ the reaction will start in opposite direction and electrochemical cell will convert the electrical energy supplied by the external cell to chemical energy which is the property of an electrolytic cell.
(D) (i) When $E_{\text{cell}} = 0$, equlibrium has been attained and concentration of Zn^{2+} and Ag^{+}becomes constant and current stops flowing.
(ii) On dilution (decrease in concentration) volume of solution, increases and number of ions per unit volume decreases. So, conductivity of a solution decreases on dilution.

OR

Given $\wedge_m = 138.9\mathrm{Scm}^2\mathrm{mol}^{-1}$ (Molar Conductivity)
Concentration $\mathrm{C} = 1.5\mathrm{mol/L}$
Conductivity $\kappa = ?$
Formula used, $\wedge_m = \frac{\kappa \times 1000}{C}$

$$\Rightarrow \quad 138.9 = \frac{\kappa \times 1000}{1.5}$$
$$\Rightarrow \quad \kappa = \frac{1.5 \times 138.9}{1000}$$
$$\Rightarrow \quad \kappa = 208.35 \times 10^{-3}\mathrm{Scm}^{-1}$$
$$\Rightarrow \quad \kappa = 2.08 \times 10^{-1}\mathrm{Scm}^{-1}$$

VSA (1-3 mark)

3. Anode: $Ni_{(s)} \rightarrow Ni^{2+}{}_{(aq)} + 2e^-$

Cathode: $Ag^+(aq) + e^- \rightarrow Ag_{(s)}$

Overall: $Ni(s) + 2Ag^+(aq) \rightarrow Ni^{2+}{}_{(aq)} + 2Ag_{(s)}$

$$E^\circ_{cell} = 0.80 - (0.25) = 0.55V, n = 2$$

$$E_{cell} = E^\circ_{cell} - \frac{0.059}{n}\log\frac{[Ni^{2+}]}{[Ag^+]^2}$$

$$= 0.55 - \frac{0.059}{2}\log\frac{[10^{-3}]}{[10^{-1}]^2}$$

$$= 0.55 - 0.0295\log(10^{-1})$$

$$= 0.55 - 0.0295(-1)$$

$$E_{cell} = 0.5795V.$$

4. $Al_{(s)}/Cd^{2+}{}_{(0.1M)}//Al^{3+}{}_{(0.01M)}/Cd_{(s)} 2Al_{(s)} + 3Cd^{2+}{}_{(0.1M)} \rightarrow 3Cd_{(s)} + 2Al^{3+}(0.01M)$

$$E_{cell} = E^\circ_{cell} - \frac{0.059}{n}\log\frac{[Al^{3+}]^2}{[Cd^{2+}]^3}$$

$$E_{cell} = 1.26 - \frac{0.059}{6}\log\frac{[0.01]^2}{[0.1]^3}$$

$$= 1.26 - \frac{0.059}{6}(-1)$$

$= 1.26 + 0.009$

$= 1.269V$

5.

Electrolytic Cell	Electrochemical Cell
(i) This is used to convert electrical energy into chemical energy	(i) This is used to convert chemical energy into electrical energy
(ii) Anode is positive and cathode is negative	(ii) Cathode is positive and anode is negative

6.

$$Zn_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Zn^{2+}{}_{(aq)} + Cu_{(s)}$$

Here, $n = 2$

$$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$$
$$= 0.34 - (-0.76) = 1.1V$$
$$F = 96500 C mol^{-1}$$
$$\Delta_r G^\circ = -nFE^\circ_{cell}$$
$$= -2 \times 1.1 \times 96500$$
$$= -212300 J mol^{-1}$$
$$= -212.3 kJ mol^{-1}$$

7. At anode: $Ni \rightarrow Ni^{2+} + 2e^-$

At cathode: $[Ag^+ + e^- \rightarrow Ag] \times 2$

Cell reaction: $Ni + 2Ag^+ \rightarrow Ni^{2+} + 2Ag$

$$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$$
$$= E^\circ Ag^+/Ag - E^\circ_{Ni^{2+}/Ni}$$
$$= 0.80V - (-0.25)V$$
$$E^\circ_{cell} = 1.05V$$
$$E^\circ_{cell} = \frac{0.0591}{n} \log K_c$$
$$\log K_c = \frac{E^\circ_{cell} \times n}{0.0591} = \frac{1.05 \times 2}{0.0591}$$
$$\log K_c = 35.53$$
$$K_c = \text{antilog } 35.53$$
$$= 3.38 \times 10^{35}$$
$$\Delta G^\circ = -nFE^\circ$$
$$= -2 \times 96500 \times 1.05$$
$$= -202650 J mol^{-1}$$
$$= -202.65 kJ mol^{-1}$$

The maximum work that can be obtained from the cell is 202.65kJ.

8. From the given cell representation,

Ag^+/Ag couple act as cathode while Zn^{2+}/Zn couple act as anode

$$\because \quad E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{Anode}$$
$$E^\circ_{cell} = 0.80 - (-0.76)$$
$$E^\circ_{cell} = 1.56V$$

The given cell reaction is,

$$Zn_{(s)} + 2Ag^+_{(aq)} \rightarrow Zn^{2+}_{(aq)} + 2Ag$$
$$n = 2$$

By Nernst equation,

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{n}\log\frac{[\text{Zn}^{2+}]}{[\text{Ag}^{+}]^{2}}$$
$$E_{\text{cell}} = E_{\text{cell}}^{\circ} + \frac{0.0591}{n}\log\frac{[\text{Ag}^{+}]^{2}}{[\text{Zn}^{2+}]}$$
$$E_{\text{cell}} = 1.56 + \frac{0.0591}{2}\log\frac{[0.01]^{2}}{[0.1]}$$
$$E_{\text{cell}} = 1.56 + \frac{0.0591}{2}\log\left(1 \times 10^{-3}\right)$$
$$E_{\text{cell}} = 1.56 - \frac{0.0591}{2} \times 3$$
$$E_{\text{cell}} = 1.56 - 0.088 = 1.472\text{V}$$

9. An electrochemical cell is a device that can generate electrical energy from the chemical reactions occurring in it.
If the external potential applied becomes greater than $\mathrm{E}_{\text{cell}}^{\circ}$ of electrochemical cell, the reaction gets reversed. It starts acting as an electrolytic cell and vice-versa.

10.

$$\mathrm{E}_{\text{cell}}^{\circ} = \mathrm{E}^{\circ}_{\text{Cd}^{+2}/\text{Cd}} - \mathrm{E}^{\circ}_{\text{Zn}^{2+}/\text{Zn}}$$
$$= -0.403\text{V} - (-0.763\text{V})$$
$$= 0.360\text{V}, n = 2$$
$$\log K_c = \left(\frac{nE_{\text{cell}}^{\circ}}{0.059}\right)$$
$$= \left(\frac{2 \times 0.360}{0.059}\right)$$
$$= \left(\frac{0.720}{0.059}\right) = 12.20$$
$$K_c = \text{antilog}\,(12.20) = 1.585 \times 10^{12}$$

11. According to Nernst equation:

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{2.303RT}{nF}\log\frac{C_2}{C_1}$$

Substituting the values $\mathrm{E}^{\circ}_{\text{cell}} = 2.71\text{V}, n = 2, \mathrm{F} = 96500\text{C}, \mathrm{C}_2 = 0.001\text{M}, \mathrm{C}_1 = 0.01\text{M}$ in eq. (i), we get,

$$E_{\text{cell}} = 2.71 - \frac{2.303 \times 8.314 \times 298}{2 \times 96500} \log \frac{0.001}{0.01}$$
$= 2.74\text{V}$

(i) When an external opposite potential applied is less than 2.71V then current will flow from Cu to Mg.

(ii) When an external opposite potential applied is more than 2.71V then current will flow from Mg to Cu.

12. $\mathrm{Sn}_{(s)}\left|\mathrm{Sn}^{2+}(0.004\mathrm{M}) \parallel \mathrm{H}^{+}(0.020\mathrm{M})\right|\mathrm{H}_{2(g)}$ (1 bar) $|\mathrm{Pt}_{(s)}$

$$\mathrm{E}^{\circ}_{\text{cell}} = \mathrm{E}^{\circ}_{(\mathrm{H}^+/\mathrm{H}_2)} - \mathrm{E}^{\circ}(\mathrm{Sn}^{2+}/\mathrm{Sn})$$
$$= 0.00 - (-0.14)$$
$$= +0.14\text{V}$$
$$\mathrm{Sn}_{(s)} \rightarrow \mathrm{Sn}^{2+}(aq) + 2\mathrm{e}^-$$
$$2\mathrm{H}^+(aq) + 2\mathrm{e}^- \rightarrow \mathrm{H}_{2(g)}$$
$$\mathrm{Sn}_{(s)} + 2\mathrm{H}^+_{(aq)} \rightarrow \mathrm{Sn}^{2+}(aq) + \mathrm{H}_{2(g)}$$
$$\mathrm{E}_{\text{cell}} = \mathrm{E}^{\circ}_{\text{cell}} - \frac{0.0591}{n} \log \frac{[\mathrm{Sn}^{2+}]}{[\mathrm{H}^+]^2}$$
$$= 0.14 - \frac{0.0591}{2} \log \frac{(4 \times 10^{-3})}{(2 \times 10^{-2})^2}$$
$$= 0.14 - 0.0295 \log 10$$
$$= 0.14 - 0.0295$$
$$= 0.1105\text{V}$$

13.

$$\Delta G^{\circ} = -nFE^{\circ}_{\text{cell}}$$
$-43600 = -2 \times 96500 \times E^{\circ}_{\text{cell}}$

$\mathrm{E}^{\circ}_{\text{cell}} = 0.23\text{V}$

$$\mathrm{E}_{\text{cell}} = \mathrm{E}^{\circ}_{\text{cell}} - \frac{0.0591}{n} \log_{10} \frac{[\mathrm{H}^+]^2[\mathrm{Cl}^-]^2}{\mathrm{P}_{\mathrm{H}^2}}$$
$$E_{\text{cell}} = 0.23\text{V} - \frac{0.0591}{2} \log_{10} \frac{[0.1]^2[0.1]^2}{1}$$
$$E_{\text{cell}} = 0.348\text{V}$$

14.

$$\mathrm{E}^{\circ}_{\text{cell}} = \mathrm{E}^{\circ}_{\mathrm{C}} - \mathrm{E}^{\circ}_{\mathrm{A}}$$
$$= -0.44 - (-0.74)$$

$$= 0.3\text{V}$$

$$E_{\text{cell}} = 0.3 - \frac{0.0591}{6}\log\frac{[\text{Cr}^{3+}]^2}{[\text{Fe}^{2+}]^3}$$

$$= 0.3 - \frac{0.0591}{6}\log\frac{(0.01)^2}{(0.1)^3}$$

$$\text{E}_{\text{cell}} = 0.309\text{V}$$

15. $\text{Al}_{(s)}|\text{Al}^{3+}{}_{(aq)}(0.01\text{M}) \parallel \text{Cu}^{2+}{}_{(aq)}(0.01\text{M})|\text{Cu}_{(s)}$

LHE: $(\text{Al(s)} \rightarrow \text{Al}^{3+}(\text{aq}) + 3\text{e}^-) \times 2$ (Oxidation at anode)

RHE: $[\text{Cu}^{2+}{}_{(aq)} + 2\text{e}^- \rightarrow \text{Cu}_{(s)}) \times 3$ (Reduction at cathode)

$$\therefore \quad n = 6$$

$$E_{\text{cell}} = \text{E}^\circ_{\text{cell}} - \frac{0.0591}{n}\log\frac{[\text{Al}^{3+}]^2}{[\text{Cu}^{2+}]^3}$$

$$\text{E}^\circ_{\text{cell}} = \text{E}_{\text{cell}} + \frac{0.0591}{n}\log\frac{[\text{Al}^{3+}]^2}{[\text{Cu}^{2+}]^3}$$

$$= 1.98 + \frac{0.0591}{6}\log\frac{(0.01)^2}{(0.01)^3}$$

$$= 1.98 + \frac{0.0591}{6}\log 10^2$$

$$= 1.98 + \frac{0.0591}{6}2\log 10$$

$$= 1.98 + \frac{0.0591}{6} \times 2 \quad [\because \log 10 = 1]$$

$$\text{E}^\circ_{\text{cell}} = 1.98 + 0.0197$$

$$\text{E}^\circ_{\text{cell}} = 1.99\text{V}$$

16. Given: $E^\circ_{\text{cell}} = 2.71\text{V}$

For the reaction,

$$\text{Mg}_{(s)} + \text{Cu}^{2+}_{(aq)} \rightarrow \text{Mg}^{2+}_{(aq)} + \text{Cu}_{(s)}$$

$$n = 2, \Delta_r\text{G}^\circ = ?$$

Using formula, $\Delta_r\text{G}^\circ = n\text{FE}_{\text{cell}}$ ell

$$\Delta_r G^\circ = -2 \times 96500\text{Cmol}^{-1} \times 2.71\text{V}$$

or

$$\Delta_r G^\circ = 523.03\text{kJmol}^{-1}$$

17. At anode:

$$\mathrm{Mg} \rightarrow \mathrm{Mg}^{2+} + 2e^{-}$$

At cathode :

Overall reaction:

$$\mathrm{Cu}^{2+} + 2\mathrm{e}^{-} \rightarrow \mathrm{Cu}$$
$$\mathrm{Mg} + \mathrm{Cu}^{2+} \rightarrow \mathrm{Mg}^{2+} + \mathrm{Cu}$$

Nernst equation:

$$\mathrm{E}_{\text{cell}} = \mathrm{E}^{\circ}_{\text{cell}} - \frac{0.059}{2} \log \frac{[\mathrm{Mg}^{2+}]}{[\mathrm{Cu}^{2+}]}$$

(at 298K)

Here,

$$E_{\text{cell}} = 2.71\mathrm{V}, [\mathrm{Mg}^{2+}] = 0.1\mathrm{M},$$
$$[\mathrm{Cu}^{2+}] = 0.01\mathrm{M}, n = 2$$
$$E_{\text{cell}} = 2.71 - \frac{0.059}{2} \log \frac{0.1}{0.01}$$
$$= 2.71 - 0.29 \times 1$$
$$= 2.681\mathrm{V}$$

18. $\frac{2}{3}\mathrm{Al_2O_3} \rightarrow \frac{4}{3}\mathrm{Al} + \mathrm{O_2}$

$$n = \frac{6 \times 2}{3} = 4e^{-}$$
$$\Delta\mathrm{G} = -n\mathrm{FE}$$
$$\Delta G = 960 \times 10^{3}\mathrm{J}, \mathrm{n} = 4$$
$$\mathrm{F} = 96500\mathrm{Cmol}^{-1}$$
$$960 \times 10^{3} = -4 \times 96500 \times E$$
$$E = -2.487V$$

Minimum potential difference needed to reduce $\mathrm{Al_2O_3} = -2.487\mathrm{V}$.

19.

$$\mathrm{A}_{(s)} + \mathrm{B}^{2+}_{(aq)} \rightleftharpoons \mathrm{A}^{2+}_{(aq)} + \mathrm{B}_{(s)}$$

Here, $n = 2$

Using formula,

$$\mathrm{E}^{\circ}_{\text{cell}} = \frac{0.059}{n} \log \mathrm{K}_c$$
$$\mathrm{E}^{\circ}_{\text{cell}} = \frac{0.059}{2} \log 10$$
$$\mathrm{E}^{\circ}_{\text{cell}} = 0.0295\mathrm{V}$$

20. To calculate the emf of the given cell, we can use the Nernst equation:

$$E_{cell} = E^\circ_{cell} - \frac{0.0591}{n} \log Q$$

Given:

- $E^\circ_{cell} = 1.05\,V$
- The reaction is: $\mathrm{Ni}(s) + 2\mathrm{Ag}^+(0.01\,M) \rightarrow \mathrm{Ni}^{2+}(0.1\,M) + 2\mathrm{Ag}(s)$
- The number of electrons transferred, $n = 2$

First, write the expression for the reaction quotient Q:

$$Q = \frac{[\mathrm{Ni}^{2+}]}{[\mathrm{Ag}^+]^2}$$

Substitute the given concentrations into Q:

$$Q = \frac{0.1}{(0.01)^2} = \frac{0.1}{0.0001} = 1000$$

Now, plug Q into the Nernst equation:

$$E_{cell} = 1.05\,V - \frac{0.0591}{2} \log 1000$$

We know $\log 1000 = 3$ (since $\log 10 = 1$):

$$E_{cell} = 1.05\,V - \frac{0.0591}{2} \times 3$$

$$E_{cell} = 1.05\,V - 0.08865\,V$$

$$E_{cell} = 0.96135\,V$$

Thus, the emf of the cell is approximately:

$$E_{cell} \approx 0.96\,V$$

2. Conductivity Of Electrolytic Solutions

21. (D) conductivity decreases whereas molar conductivity increases on dilution.

22. (b) 0.1MHCl solution

Conductivity is higher for strong electrolyte, conductivity decreases with dilution.

23. (a) 0.01M

Explanation: Since, $\Lambda_m = \frac{\kappa \times 1000}{M}$

o, $\Lambda_m \propto \frac{1}{M}$
Hence, the value of concentration (molarity), is inversely proportional to the value of molar conductivity.
So, out of the given four solutions, $0.01M$ is the lowest concentration, so $0.01MKCl$ solution will have the highest value of molar conductivity.

24. (a) Both (A) and (R) are true and (R) is the correct explanation of (A).
Explanation: Conductivity changes with the concentration of the electrolyte. It always decreases with the decrease in concentration, for both, weak and strong electrolytes. This can be explained by the fact that the number of ions per unit volume that carry the current in a solution decreases on dilution.

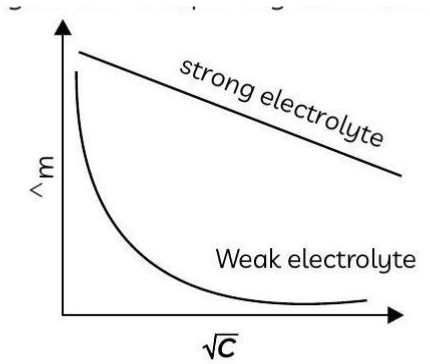
25. (c) $390.5scm^2mol^{-1}$
Explanation : Limiting molar conductivity

$$\begin{aligned}\lambda^{\circ}_{mCH_3COOH} &= \lambda^{\circ}_m CH_3COO^- + \lambda^{\circ}_m H^+ \\ &= \lambda^{\circ}_{mCH_3COONa} - \lambda^{\circ}_{mNa^+} \\ &= 91 - 50.1 + 349.6 \\ &= 390.5Scm^{\circ}_m mol^{-1}\end{aligned}$$

26. D) 271.6 S cm^2 mol^{-1}

VSA (1-3 mark)

27. ' B ' is a strong electrolyte. The molar conductivity increases slowly with dilution as there is no increase in number of ions on dilution as strong electrolytes are completely dissociated.



28. (i) A-Strong electrolyte, B-Weak electrolyte
(ii) Λ°_m for weak electrolytes cannot be obtained by extrapolation while Λ°_m for strong electrolytes can be obtained by intercept.

29. Conductivity of CH_3COOH decrease with dilution because the number of ions per cm^3 of the solution decreases on the dilution of the ions thus the current carried by them also decreases due to which the current passing through the solution decreases.

30. Conductivity is directly proportional to the number of ions present in unit volume of solution. When a solution of acetic acid is diluted, the degree of dissociation of acetic acid increases and the total number of ions increases, but the number of ions per unit volume decreases. Hence, conductivity of CH_3COOH decreases on dilution.

31. The limiting molar conductivity of an electrolyte is defined as its molar conductivity when the concentration of the electrolyte in the solution approaches to zero.
The conductivity of an electrolyte solution is the conductance of ions present in a unit volume of the solution. The number of ions (responsible for carrying current) decreases when the solution is diluted or the concentration is decreased. As a result, the conductivity of an electrolyte solution decreases with the decrease in concentration.

32. The fraction of the total number of the molecules dissociate in solution as ions is known as the degree of dissociation.

$$\text{Molar conductivity } (\lambda_m) = \alpha\lambda_m^\circ$$

Where λ_m° is the molar conductivity at the infinite dilution.

33. The reciprocal of the resistivity is known as conductivity. It is denoted by Kappa.

$$\kappa = 1/R$$

Where, $\kappa =$ conductivity

$$\text{R} = \text{ resistivity}$$

Molar conductivity is defined as the conducting power of all the ions produced by dissolving one mole of an electrolyte in solution.

$$\lambda_m = \frac{K}{C}$$

Where $\lambda_m =$ Molar conductivity

$$\text{K} = \text{ constant}$$
$$\text{C} = \text{ Concentration}$$

The conductivity of the solution decrease with the decrease in the concentration of both the weak and the strong electrolyte because the number of ions per unit volume of the solution decreases and the current is also carried by these ions which results in the conductance of the solution

thus the conductance of the solution decreases whereas the molar conductivity increase with the decrease in the concentration of the ions since the volume of the solution increases due to which the total number of the current carrying ions also increases. Thus, the dilution is more compensated by the increase in the volume of the solution.

34. Degree of dissociation can be defined as the ratio of molar conductivity at a specific concentration to the molar conductivity at infinite dilution.
Degree of dissociation $(\alpha) = \frac{\Lambda_m^c}{\Lambda_m^\infty}$

35. Kohlrausch's law of independent migration of ions states that at infinite dilution, the limiting molar conductivity of an electrolyte can be represented as the sum of individual contributions of its cations and anions. The conductivity of a solution decreases with dilution because less ions are present for conduction.

36. Conductivity is the conductance of a solution of 1cm length and having 1sq. cm as the area of cross section whereas molar conductivity is defined as the conductance of all the ions obtained by dissolving 1 mole of an electrolyte. As the kinetic speed of an ion increases with temperature. Thus, conductivity and molar conductivity both increases.

37.

1. Calculate Molar Conductivity (Λ_m):
$\Lambda_m = \frac{\kappa}{C}$
Where:

- $\kappa = 2.48 \times 10^{-2}\,\mathrm{S\ cm^{-1}}$
- $C = 0.2\,\mathrm{M}$

$\Lambda_m = \frac{2.48\times10^{-2}}{0.2} = 0.124\,\mathrm{S\ cm^2mol^{-1}}$

2. Calculate Limiting Molar Conductivity (Λ_m^0):
$\Lambda_m^0 = \lambda_0^{\mathrm{K^+}} + \lambda_0^{\mathrm{Cl^-}}$
$\Lambda_m^0 = 73.5 + 76.5 = 150\,\mathrm{S\ cm^2mol^{-1}}$

3. Calculate Degree of Dissociation (α):

$$\alpha = \frac{\Lambda_m}{\Lambda_m^0}$$
$$\alpha = \frac{0.124}{150}$$
$$\alpha = 0.0008267 \approx 8.27 \times 10^{-4}$$

The molar conductivity of the solution is:

$$\boxed{0.124\ \mathrm{S\ cm^2 mol^{-1}}}$$

The degree of dissociation is:

$$\boxed{8.27 \times 10^{-4}}$$

38.

1. **Calculate the cell constant (G*):**

$$\text{Cell constant (G*)} = \text{Conductivity} \times \text{Resistance}$$
$$\mathrm{G^*} = 0.0248\ \mathrm{S\ cm^{-1}} \times 200\ \Omega$$
$$\mathrm{G^*} = 4.96\ \mathrm{cm^{-1}}$$

2. **Calculate the conductivity (κ) of 0.05 mol L^{-1} KCl solution:**

$$\text{Conductivity}(\kappa) = \frac{\text{Cell constant (G*)}}{\text{Resistance}}$$
$$\kappa = \frac{4.96\ \mathrm{cm^{-1}}}{620\ \Omega}$$
$$\kappa = 0.008\ \mathrm{S\ cm^{-1}}$$

3. **Calculate the molar conductivity (Λ_m) of 0.05 mol L^{-1} KCl solution:**

$$\Lambda_m = \frac{\kappa \times 1000}{c}$$

where c is the concentration in mol L^{-1}.

$$\Lambda_m = \frac{0.008\ \mathrm{S\ cm^{-1}} \times 1000}{0.05\ \mathrm{mol\ L^{-1}}}$$
$$\Lambda_m = \frac{8\ \mathrm{S\ cm^{-1}}}{0.05}$$
$$\Lambda_m = 160\ \mathrm{S\ cm^2 mol^{-1}}$$

Summary:

- The conductivity of 0.05 mol L^{-1} KCl solution is $0.008\ \mathrm{S\ cm^{-1}}$.
- The molar conductivity of 0.05 mol L^{-1} KCl solution is $160\ \mathrm{S\ cm^2 mol^{-1}}$.

Long (4-5 mark)

39. (A) Yes, if the concentration of $\mathrm{ZnSO_4}$ in the two half cell is different, the electrode potential will be different making the cell possible.

(B) $\lambda^0{}_m(\mathrm{MgCl_2}) = \lambda^0{}_\mathrm{m}(\mathrm{Mg^{2+}}) + 2\lambda^0{}_\mathrm{m}(\mathrm{Cl^-})$

$$258.6 = 106 + 2\lambda^0_m(\mathrm{Cl^-})$$
$$\lambda^0{}_\mathrm{m}(\mathrm{Cl^-}) = 76.3\mathrm{Scm^2mol^{-1}}$$

(C) Cell constant $\mathrm{G}^* = k \times \mathrm{R}$

$$k = \frac{G^*}{R} = \frac{0.146}{1000}$$
$$= 1.46 \times 10^{-4}\mathrm{Scm^{-1}}.$$

40. (A)

$$\Lambda_m = \frac{\kappa \times 1000}{M}$$
$$= \frac{8 \times 10^{-5} \times 10^3}{2 \times 10^{-3}}$$
$$= 40\mathrm{scm^2mol^{-1}}$$
$$\alpha = \frac{\Lambda_m}{\Lambda_m^\circ} = \frac{40}{404}$$
$$\alpha = 0.099$$

(B) $\mathrm{Ni}_{(s)} + 2\mathrm{Ag}^+_{(aq)} \rightleftharpoons \mathrm{Ni}^{2+}{}_{(aq)} + 2\mathrm{Ag}_{(s)}$

$$n = 2$$

$\mathrm{E}^\circ_\mathrm{cell} = \mathrm{E}^\circ_\mathrm{Ag}/\mathrm{Ag}^+ - \mathrm{E}^\circ\mathrm{Ni}^{2+}/\mathrm{Ni}$

$$= 0.80 - (-0.25)$$
$$= 0.80 + 0.25$$
$$= 1.05\mathrm{V}$$
$$\Delta\mathrm{G}^\circ = -nF\mathrm{E}^\circ{}_\mathrm{cell}$$
$$= -2 \times 96500 \times 1.05$$
$$= 202650\mathrm{J}$$

Now, $\Delta G^\circ = -2.303RT\log K_c$

$$\log K_c = \frac{-\Delta G^\circ}{2.303RT}$$
$$= \frac{202650}{2.303 \times 8.314 \times 298}$$

$\log K_c = 35.516$

41. (A) The cell potential remains constant during its lifetime as the overall reaction does not involve any ion in solution whose concentration can change during its lifetime.

(B) $KCl_{(aq)} \rightarrow K^+_{(aq)} + Cl_{(aq)}$
Cathode: $H_2O_{(b)} + e^- \rightarrow 1/2H_{2(g)} + OH^-_{(aq)}$
Anode: $Cl^-_{(aq)} \rightarrow 1/2Cl_{2(aq)} + e^-$
Net reaction:
$KCl_{(aq)} + H_2O_{(b)} \rightarrow K^+_{(aq)} + OH^-_{(aq)} + 1/2H_{2(g)}$
$+1/2Cl_{2(g)}$
(C) Given, potential of hydrogen gas electrode $=- 0.59V$
Electrode reaction: $H^+ + e^- \rightarrow 0.5H_2$
Applying Nernst equation,

$$E\left(\frac{H^+}{H_2}\right) = E^\circ\left(\frac{H^+}{H_2}\right) - \frac{0.059}{n}\log\frac{[H_2]^{1/2}}{[H^+]}$$

$$\left(\frac{H^+}{H_2}\right) = 0V$$

$$\left(\frac{H^+}{H_2}\right)_{n=1} =- 0.59V$$

$$[H_2]= 1 \text{ bar}$$

$$-0.59= 0 - 0.059\left(-\log [H^+]\right)$$

$$-0.59=- 0.059pH$$

$$\therefore \quad pH= 10$$

42. (A) "A" is copper, metals are conductors thus have high value of conductivity.
(B) $Mg^{2+} + 2e^- \rightarrow Mg$
1 mole of magnesium ions gains two moles of electrons or $2F$ to form 1 mole of Mg
$24gMg$ requires $2F$ electricity $4.8gMg$ requires $2 \times 4.8/24 = 0.4F$
$= 0.4 \times 96500 = 38600C$
$Ca^{2+} + 2e^- \rightarrow Ca$
$2F$ electricity is required to produce 1 mole $= 40gCa$
$0.4F$ electricity will produce $8gCa$
(C) $F = 96500C, n = 2,$
$Sn^{2+}{}_{(aq)} + 2e^- \rightarrow Sn_{(s)} E^\circ = 0.14V$
$Cu^{2+}{}_{(aq)} + e^- \rightarrow Cu^+{}_{(aq)}, E^\circ = 0.15V$
$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$
$= 0.15 - (- 0.14) = 0.29V$
$\Delta G^\circ =- nnE^\circ{}_{cell}$
$=- 2 \times 96500 \times 0.29 =- 55970J/mol$
As $\Delta G =-$ ve
The reaction is spontaneous in nature.

43. (A) $\Lambda^\circ_{m(\mathrm{CH_3COOH})} = \Lambda^\circ\mathrm{CH_3COO^-} + \Lambda^\circ\mathrm{H^+}$

$$= \Lambda^\circ\mathrm{CH_3COO^-} + \Lambda^\circ\mathrm{Na^+} + \Lambda^\circ_{\mathrm{H}} + \Lambda^\circ\mathrm{Cl^-}$$
$$- (\Lambda_{\mathrm{Na^+}} + \Lambda^\circ\mathrm{Cl^-})$$
$$= \Lambda^\circ_{m(\mathrm{CH_2COONa})} + \Lambda^\circ_{m(\mathrm{HCl})} - \Lambda^\circ_{m(\mathrm{NaCl})}$$
$$= (91.0 + 425.9 - 126.4)\mathrm{Scm^2mol^{-1}}$$
$$= 390.5\mathrm{Scm^2mol^{-1}}$$

(B) $\alpha = \frac{\Lambda_m}{\Lambda^\circ_m}$

$$\alpha = \frac{39.0}{390.5}$$
$$= 0.1$$

% ionization $= \alpha \times 100\%$

$$= 10\%$$

% unionized $= 90\%$

44. (A) $\Lambda^\circ_{(\mathrm{CH_3COOH})} = \lambda^\circ_{\mathrm{H^+}} + \lambda^\circ_{\mathrm{CH_3COO^-}}$

$$= 349.6 + 40.9$$
$$= 390.5\mathrm{Scm^2mol^{-1}}$$

$$\Lambda_m = \frac{\kappa \times 100}{c}$$

$$\Lambda_m = \frac{8.0 \times 10^{-5}\mathrm{Scm^{-1}} \times 1000\mathrm{cm^3L^{-1}}}{0.0024\mathrm{molL^{-1}}}$$
$$= 33.33\mathrm{Scm^2mol^{-1}}$$

$$\alpha = \frac{\Lambda_m}{\Lambda^\circ_m}$$

$$\alpha = \frac{33.33\mathrm{Scm^2mol^{-1}}}{390.5\mathrm{Scm^2mol^{-1}}}$$
$$= 0.085$$

(B) Electrolyte B is a strong electrlyte.
For a strong electrolyte, limiting molar conductivity increases only to a lesser extent because interionic interactions are overcome during dilution. Limiting molar conductivity increases to a greater extent for a weak electrolyte when the degree of dissociation increases with dilution, resulting in an increase in the number of ions in the total volume of solution.

45.

$$\alpha = \frac{\Lambda_\mathrm{m}}{\Lambda^\circ_\mathrm{m}}$$

$$\mathring{A}_{\mathrm{m(CH_3COOH)}} = \lambda^\circ_{(\mathrm{CH_3COO^-})} + \lambda^\circ_{(\mathrm{H^+})}$$
$$= 40.9 + 349.6$$
$$= 390.5\mathrm{Scm^2mol^{-1}}$$
$$\therefore \quad = 39.05\mathrm{Scm^2mol^{-1}}$$
$$\therefore \quad \alpha = \frac{39.05}{390.5} = 0.1$$

Numerical Type Qs (1 - 3 marks)

46. Given: Conductivity, $\kappa = 0.0248 \mathrm{Scm}^{-1}$
Molarity, $C = 0.20\mathrm{M} = 0.20 \mathrm{molL}^{-1}$
Using formula, $\Lambda_m^c = \frac{1000 \times \kappa}{C}$

$$\Lambda_m^c = \frac{1000 \times 0.0248}{0.20}$$
$$= 124 \mathrm{Scm}^2\mathrm{mol}^{-1}$$
$$\alpha = \frac{\Lambda_m^c}{\Lambda_m^\circ} = \frac{124}{73.5 + 76.5} = 0.82$$

47. Resistance of $0.1\mathrm{MKCl}$ solution $R = 100\Omega$
Conductivity $\kappa = 1.29 \mathrm{Sm}^{-1}$
Cell constant $\mathrm{G}^* = \kappa \times R = 1.29 \times 100 = 129 \mathrm{m}^{-1}$
Resistance of $0.02\mathrm{MKCl}$ solution, $R = 520\Omega$
Conductivity,

$$\kappa = \frac{\text{cell constant}}{R}$$
$$= \frac{129 \mathrm{m}^{-1}}{520\Omega}$$
$$= 0.248 \mathrm{Sm}^{-1}$$
$$C = 0.02 \mathrm{molL}^{-1}$$
$$= 1000 \times 0.02 \mathrm{molm}^{-3}$$
$$= 20 \mathrm{molm}^{-3}$$

Concentration,
Molar conductivity,

$$\Lambda_m = \frac{\kappa}{C}$$
$$= \frac{0.248 \mathrm{Sm}^{-1}}{20 \mathrm{molm}^{-3}}$$
$$= 0.0124 \mathrm{Sm}^2\mathrm{mol}^{-1}$$

4. Electrolytic Cell And Electrolysis

MCQ

48. (B) H_2–O_2 fuel cell

VSA (1-3 mark)

49. To convert 1mol of H_2O to O_2 following equation is used:

$$2H_2O_2 \longrightarrow 2H_2O + O_2$$

It is evident from the above equation that 2 moles of H_2O are required to convert to O_2.

1 mole has 96500 Coulombs thus 2 moles will have:

$$2 \times 96500 = 193000 \text{ Coulombs}$$

Thus, 193000 Coulombs are required for the conversion of 1mol of H_2O to O_2.

50. The reaction at anode with lower value E° is preferred, i.e., O_2 gas should be liberated on account of overpotential of oxygen reaction anode, preferred reaction is:

$$Cl^-_{(aq)} \rightarrow \frac{1}{2}Cl_{2(g)} + e^-$$

i.e., Cl_2 gas is liberated at anode in the electron of an $NaCl$.

51.

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 5H_2O$$

1 mole of MnO_4^- requires 5 mole of electrons = 5 Faraday.

52. Oxidation of H_2O to O_2 is:

$$2H_2O \rightarrow 4H^+ + O_2 + 4e^-$$

Oxidation of 2 moles of water requires $96500 \times 4C$

∴ Oxidation of 1 mole of water requires

$$= \frac{96500 \times 4}{2}$$

$$= 1.93 \times 10^5 C$$

53. Higher the standard reduction potential of a species, more easily it is reduced at the cathode. As $Ag^+(aq)$ has greater standard reduction potential, therefore, the reaction that will occur at the cathode is:

$$Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$$

54. $Zn^{2+}_{(aq)} + 2e^- \longrightarrow Zn_{(s)}$

$1\text{mol} 2\text{mol}$

One mole of Zn^{2+} requires 2 moles of electron reduction, i.e.,

$Q = 2 \times F = 2 \times 96500 = 193000C$

55. Faraday's-First Law of Electrolysis: It is one of the primary laws of electrolysis. It states, during electrolysis, the amount of chemical reaction which occurs at any electrode under the influence of electrical energy is proportional to the quantity of electricity passed through the electrolyte.

56. Higher the standard reduction potential of a species, more easily it is reduced at the cathode. As $Cu^{2+}{}_{(aq)}$ has greater standard reduction potential, therefore, the reaction that will occur at the cathode is:

$$Cu^{2+}{}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$$

57. (a) Determine the cell reaction and calculate the standard emf:
The cell reaction can be written as:

$$\mathrm{Sn}(s) + 2\mathrm{H}^+(aq) \rightarrow \mathrm{Sn}^{2+}(aq) + \mathrm{H}_2(g)$$

Using the given standard electrode potentials:

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

$$E^\circ_{\text{cell}} = 0.00\ \mathrm{V} - (-0.14\ \mathrm{V}) = 0.14\ \mathrm{V}$$

(b) Calculate the reaction quotient Q:

$$Q = \frac{[\mathrm{Sn}^{2+}]}{[\mathrm{H}^+]^2}$$

Given concentrations:

$$[\mathrm{Sn}^{2+}] = 0.001\ \mathrm{M},\ [\mathrm{H}^+] = 0.01\ \mathrm{M}$$

$$Q = \frac{0.001}{(0.01)^2} = 0.001/0.0001 = 10$$

(c) Apply the Nernst equation:

$$E_{\text{cell}} = 0.14\ \mathrm{V} - \frac{0.0591}{2}\log 10$$

$$E_{\text{cell}} = 0.14\ \mathrm{V} - \frac{0.0591}{2} \times 1$$

Numerical Type Qs (1 - 3 marks)

58. Charge $= 24,000$ Coulombs

According to reaction,

$$Cr^{6+} + 6e^{-} \rightarrow Cr_{(s)}$$

We require 6F or 6×96500C to deposit 1 mole or 52g of Cr

For 24,000 C, the mass of Cr deposited

$$\frac{52 \times 24,000}{6 \times 96500} = 2.15\text{g}$$

According to Faraday's second law of electrolysis,

$$\frac{W_1}{W_2} = \frac{Q_1}{Q_2}$$

$$\text{or} \quad \frac{2.15}{1.5} = \frac{24,000}{12.5 \times t}$$

$$t = \frac{24,000 \times 1.5}{12.5 \times 2.15} = 1340\text{s}$$

59. According to Faraday's first law of electrolysis, mass of substance deposited ∝ quantity of electricity passed

$$\Rightarrow W = ZQ = \frac{E}{F} \times Q$$

$$\Rightarrow 2.8 = \frac{28}{96500} \times Q$$

$$\Rightarrow Q$$

$$\Rightarrow = \frac{2.8 \times 96500}{28} = 9650\text{C}$$

Now

$$Q = |t$$

$$\Rightarrow \quad t = \frac{Q}{1} = \frac{9650}{2} = 4825\text{s}$$

According to Faraday's second law of electrolysis,

$$\frac{\text{Mass of Fe deposited}}{\text{Mass of } Zn \text{ deposited}} = \frac{\text{Eq. wt. of Fe}}{\text{Eq. wt. of } Zn}$$

$$\Rightarrow \quad \frac{2.8}{\text{Mass of } Zn \text{ deposited}} = \frac{28}{32.65}$$

$$\Rightarrow \quad \text{Mass of } Zn \text{ deposited} = \frac{2.8}{28} \times 32.65$$

$$= 3.265\text{g}$$

60. Given $\text{I} = 2\text{A}, t = 15\text{min} = 15 \times 60\text{s} = 900sw = ?$

$Q = 1 \times t = 2 \times 900 = 1800\text{C}$

Reaction for deposition of Ag is as follows:

$$\begin{array}{ll} Ag^{+} + 1e^{-} & \rightarrow Ag \\ 1F & 1mol \\ 1 \times 96500C & 108g \end{array}$$

Thus, 1×96500C of electricity is required to deposit 108g of Ag.

∴ 1800 C of electricity would deposit

$$= \frac{108 \times 1800}{1 \times 96500}$$
$$= 2.014\text{g of Ag}$$

61. (A) $2Fe^{3+} + 2e^{-} \rightarrow 2Fe^{2+}$

and $2I^{-} \rightarrow I_2 + 2e^{-}$

Hence, for the given cell reaction, $n = 2$

$\Delta_r G^{\circ} =- nFE^{\circ}_{\text{cell}} =- 2 \times 96500 \times 0.236$

$$=- 45.55\text{kJmol}^{-1}$$
$$\Delta_r G^{\circ} =- 2.303 RT \log K_c$$

or $\log K_c = \frac{\Delta_r G^{\circ}}{2.303RT}$

$$= \frac{-45.55\text{kJmol}^{-1}}{2.303 \times 8.314 \times 10^{-3}\text{kJK}^{-1}\text{mol}^{-1} \times 298\text{K}}$$

∴ 7.94

$K_c =$

$= 9.616 \times 10^{7}$

(B) Given, I = 0.5A, $t = 2$hrs.

Number of electrons = ?

Total charge $(Q) = 1 \times t$

$$= 0.5 \times 2 \times 60 \times 60$$
$$= 3600\text{C}$$

∴ Number of electrons

$$= \frac{\text{Total charge}}{\text{Charge of one electron}}$$
$$= \frac{3600}{1.6 \times 10^{-19}}$$
$$= 2.25 \times 10^{22} \text{ electrons}$$

62. Faraday's First law states that: "It is one of the primary laws of electrolysis. It states that during electrolysis, the amount of chemical reaction which occurs at any electrode under the influence of electrical energy is proportional to the quantity of electricity passed through the electrolyte".

$$Cu^{2+} + 2e^{-} \rightarrow Cu$$

63. Given:

$$\text{Current} = 5\text{A}$$
$$\text{Time} = 20 \times 60 = 1200\text{s}$$
$$\therefore \quad \text{Charge} = \text{Current} \times \text{Time}$$
$$= 5 \times 1200$$
$$= 6000\text{C}$$

According to the reaction,

$$\text{Ni}^{2+}{}_{(aq)} + 2e^- \rightarrow \underset{58.7\text{g}}{\text{Ni}_{(s)}}$$

Nickel deposited by $2 \times 96487\text{C} = 58.71\text{g}$

Therefore, nickel deposited by 6000C

$$= \frac{58.71 \times 6000}{2 \times 96487}\text{g}$$
$$= 1.825\text{g}$$

Hence, 1.825g of nickel will be deposited at the cathode.

4. Commercial Cells/ Batteries

Case Based Qs (4-5 mark)

64. (A) $2\text{mol}e^-$(or 2F) have been transferred from anode to cathode to consume 2mol of H_2SO_4 therefore, one mole H_2SO_4 requires one faraday of electricity or 96500 coulombs.

(B) $W_{\text{max}} =- nFE^\circ =- 2 \times 96500 \times 2.0 = 386000\text{J}$ of work can be extracted using lead storage cell when the cell is in use.

(C) Both yes and no should be accepted as correct answers depending upon what explanation is provided.

Yes, Hydrogen is a fuel that on combustion gives water as a byproduct. There are no carbon emissions and no pollutions caused. However, at present the means to obtain hydrogen are electrolysis of water which use electricity obtained from fossil fuels and increase carbon emissions.

Inspite of the problems faced today in the extraction of hydrogen, we cannot disagree on the fact that hydrogen is a clean source of energy. Further research can help in finding solutions and greens ways like using solar energy for extraction of hydrogen.

No. It is true that hydrogen is a fuel that on combustion gives water as a byproduct. There are no carbon emissions and no pollutions caused. However, at present the means to obtain hydrogen are electrolysis of water which use electricity obtained from fossil fuels and increase carbon emissions. Hydrogen is no doubt a green fuel, but the process of extraction is not green as of today. At present, looking at the process of extraction, hydrogen is not a green fuel.

OR

(Both answers will be treated as correct)

(i) Lead batteries are currently the most important and widely used batteries. These are rechargeable. The problem is waste management which needs research and awareness. Currently, these are being thrown into landfills and there is no safe method of disposal or recycling. Research into safer method of disposal will reduce the pollution and health hazards caused to a great extent.

(ii) Fuel cell is a clean source of energy. Hydrogen undergoes combustion to produce water. The need of the hour is green fuel and hydrogen is a clean fuel. The current problem is obtaining hydrogen. Research that goes into this area will help solve the problem of pollution and will be a sustainable solution.

VSA (1-3 mark)

65. (A) Mercury cell delivers a constant potential of approx $1.35\mathrm{V}$ during its life-time because the overall cell reaction does not involve any ion in solution whose concentration can change during its life-time.

Overall reaction is:

$\mathrm{Zn}_{(s)} + \mathrm{HgO}_{(s)} \rightarrow \mathrm{ZnO}_{(s)} + \mathrm{Hg}_{(1)}$

(B) In the experimental determination of electrolytic conductance, direct current (DC) is not used because it changes the composition of the solution.

OR

Galvanic cells that are designed to convert the energy of combustion of fuels like hydrogen, methane, methanol, etc., directly into electrical energy are called fuel cells. For example, $\mathrm{H}_2 - \mathrm{O}_2$ fuel cell (i.e., hydrogen-oxygen fuel cell)

In fuel cells, reactants are fed continuously to the electrodes and products are removed continuously from the electrolyte compartment while primary batteries contain a limited amount of reactants and are discharged when the reactants are consumed completely and secondary batteries can be recharged but it takes a long time.

Thus, the fuel cells have advantages over primary and secondary batteries.

66. Anode:

$$\mathrm{Zn}_{(s)} \rightarrow \mathrm{Zn}^{2+}{}_{(aq)} + 2e^-$$

Cathode:

$$\mathrm{O}_{2(g)} + 2\mathrm{H}_2\mathrm{O}_{(n)} + 4e^- \rightarrow 4\mathrm{OH}^-_{(aq)}.$$

Overall:

$$2Zn_{(s)} + O_{2(g)} + 2H_2O_{(l)} \rightarrow 2Zn^{2+}(aq)$$
$$2Zn_{(s)} + O_{2(g)} + 2H_2OH^-_{(l)} \rightarrow 2Zn(Oq).$$
$$(OH)_{2(ppt)}$$

67. The hydrogen oxygen fuel cell was used in the Apollo space programme.

68. The advantages of the fuel cells are:
(1) They are pollution free.
(2) Their efficiency is around 70% compared with that of the thermal plants of around 40%.

69. The galvanic cells which gives the direct electrical energy by the combustion of fuels like hydrogen, methane and methanol etc, are called fuel cells.
Working: The working of fuel cell involve passing of hydrogen and oxygen into a concentrated solution of sodium hydroxide via carbon electrodes.
Cathode: $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$
Anode: $2H_2 + 4OH^- \rightarrow 4H_2O + 4e^-$.

70. Cells that convert the energy of combustion of fuels (like hydrogen, methane, methanol, etc.) directly into electrical energy are called fuel cells.
Advantages:
(i) Its efficiency is about 75% which is considerably higher than conventional cells.
(ii) It is non-polluting in nature as no harmful products are formed.

71. The mercury cell is generally used in heating aids.
Reaction at cathode:

$$HgO + H_2O + 2e^- \rightarrow Hg + 2OH^+$$

Reaction at anode:

$$Zn_{(s)} + 2OH^- \rightarrow ZnO_{(s)} + H_2O + 2e^-$$

72. The cell which is generally used in inverters is secondary cell, i.e., lead storage battery.

At cathode: $PbO_{2(s)} + SO^{2-}_{4(aq)} + 4H^+_{(aq)} + 2e^-$
$\rightarrow PbSO_{4(s)} + 2H_2O$

At anode: $Pb_{(s)} + SO^{2-}_{4(aq)} \rightarrow PbSO_{4(s)} + 2e^-$

73. Metals of lower electrode potential value when connected with iron protect it from oxidation and prevents the corrosion. Hence, coating of metal A having the lower electrode
potential will be better than metal B which has the higher electrode potential.

74. The type of battery which can be recharged again and again by passing the current through it in the opposite direction so that it can be used again. For example: Lead storage battery.A mercury battery (also called mercuric oxide battery) is a non-rechargeable electrochemical battery and a primary cell.

75. Mercury cell is more advantageous than dry cell because dry cell has a very short life span due to the conversion of zinc to zinc chloride that makes the zinc casing porous. Due to this porous casing, the substance inside the cell leaks out and corrodes the metal, reducing the lifetime of the cell. While, in case of mercury cell, the overall reaction does not involve formation of any ion in the solution whose concentration can change during its life time.

Click Here to Download More Chapters